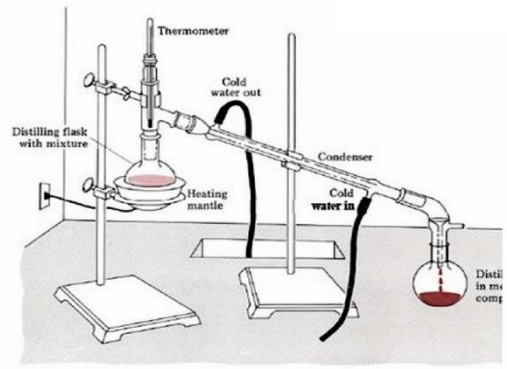

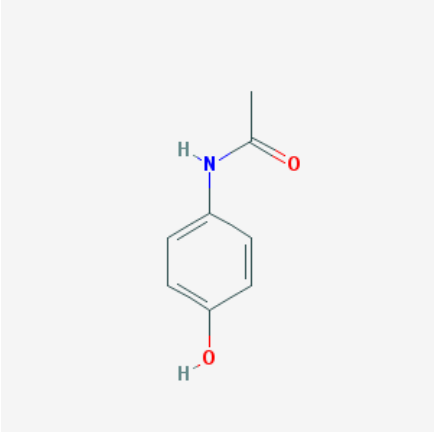

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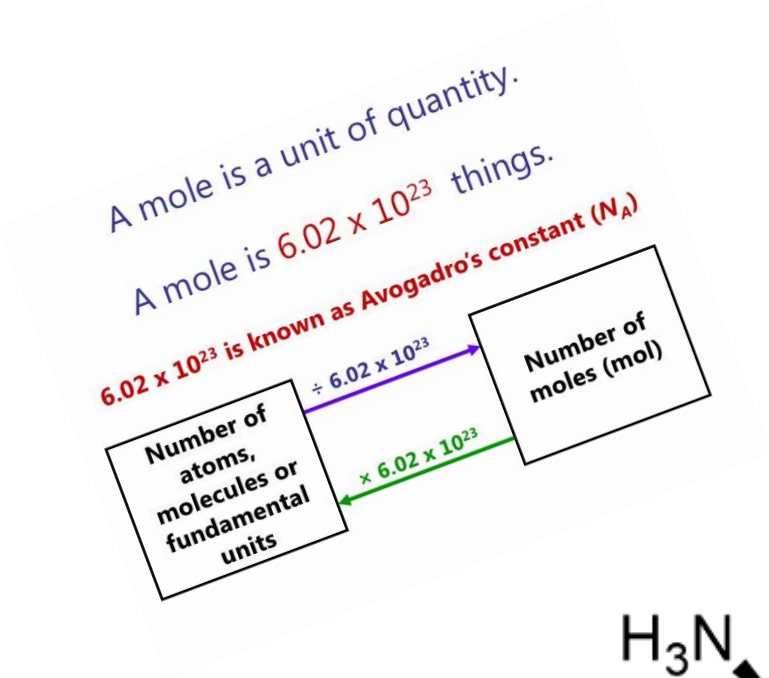

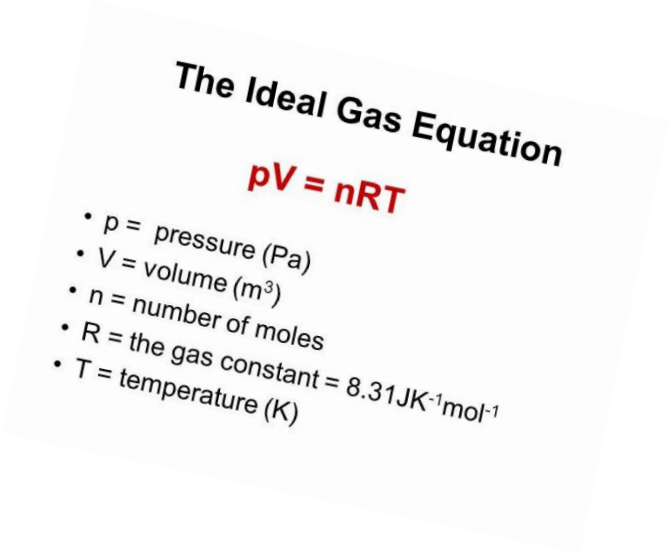

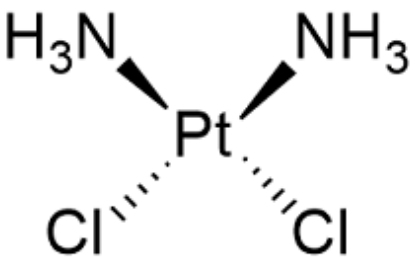

Cisplatin

AQA A level Chemistry Transition work

Name ..... **Date** .....

Aims

This worksheet gives you practice in converting mass into moles and vice versa. It also gives you practice in calculating the masses of reactants (or products) using a balanced symbol equation and the mass of another reactant (or product).

Learning outcomes

After completing this worksheet, you should be able to:

- calculate the relative formula mass for one substance when the relative formula masses are given for all the other substances in a balanced symbol equation
- interpret balanced symbol equations in terms of mole ratios
- use balanced symbol equations to calculate reacting masses.

Setting the scene

The reaction between H_2 and O_2 to make water, H_2O, is:

$$2H_2 + O_2 \rightarrow 2H_2O$$

From the balanced symbol equation, we know that two molecules of hydrogen, H_2, react with one molecule of oxygen, O_2, to form two molecules of water, H_2O. However, when working with chemicals it is very difficult to measure the mass of an individual atom or molecule because these masses are so tiny. Instead, chemists have invented a unit that is similar to the way we might say that one dozen eggs contains 12 eggs. This unit is called a mole: 1 mole of an element contains 6.02×10^{23} atoms and 1 mole of a compound contains 6.02×10^{23} molecules or ions.

The relative atomic mass (A_r) in grams of an element contains 1 mole of atoms. The relative formula mass (M_r) in grams of a compound contains 1 mole of molecules or ions. Therefore, if we know the mass of an element or compound in grams, we can calculate the number of moles it contains using the formula:

$$\text{number of moles (mol)} = \frac{\text{mass (g)}}{A_r} \text{ or } \text{number of moles (mol)} = \frac{\text{mass (g)}}{M_r}$$

We can also use the number of moles involved in a chemical equation to calculate the number of moles – and therefore the masses – of all the other reactants or products in that equation.

Worked examples

Example 1

Calculate the number of moles in 45 g of water, H_2O.

(A_r values: O = 16, H = 1)

Step 1

First, calculate the M_r of water:

$= (2 \times H) + (1 \times O)$

$= (2 \times 1) + (1 \times 16)$

$= 18$

Step 2

Then use the equation:

$$\text{number of moles} = \frac{\text{mass (g)}}{M_r}$$

to calculate the amount in moles:

$$\text{number of moles} = \frac{45\text{ g}}{18} = 2.5\text{ mol}$$

So 45 g of water contains 2.5 moles.

Example 2

Calculate the mass of oxygen that must react to make 45 g of water.

(A_r values: O = 16, H = 1)

Step 1

First write out the balanced symbol equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

Step 2

Then calculate the number of moles in 45 g of water – we know from Example 1 that the answer is 2.5 moles.

Step 3

Next, use the ratio of moles given in the balanced equation to calculate the number of moles of oxygen that must react to make 2.5 moles of water.

From the balanced equation, you can see that 2 moles of hydrogen react with 1 mole of oxygen to make 2 moles of water.

So the oxygen and the water appear in the balanced symbol equation in a 1 : 2 ratio. Therefore, the number of moles of oxygen that must react = 2.5 mol ÷ 2 = 1.25 mol.

Step 4

Finally, calculate the mass of oxygen that contains 1.25 moles:

$$\text{number of moles (mol)} = \frac{\text{mass (g)}}{M_r} \text{ therefore}$$

$$\text{mass (g)} = \text{number of moles (mol)} \times M_r$$

M_r O_2 = (2 × O) = (2 × 16) = 32

Substituting into the equation:

mass (g) = 1.25 mol × 32 = 40

So 40 g of oxygen is needed to produce 45 g of water.

Questions

Relative atomic masses (A_r): H = 1; C = 12; N = 14; O = 16; Mg = 24; Ca = 40

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Name ..... **Date** .....

1 Calculate the mass of one mole of each of the following substances:

a Ammonia, NH_3

b Methane, CH_4

c Calcium carbonate, $CaCO_3$

d Magnesium hydroxide, $Mg(OH)_2$. (*4 marks*)

2 **a** Calculate the mass of the following:

i 3 moles of ammonia

ii 4.5 moles of methane

iii 0.2 moles of calcium carbonate

iv 1.4 moles of magnesium hydroxide. (*4 marks*)

b Calculate the number of moles in the following:

i 136 g of ammonia

ii 160 g of methane

iii 250 g of calcium carbonate

iv 2.9 g of magnesium hydroxide. (*4 marks*)

c 15.3 g of an unknown substance was found to contain 0.15 moles.

Calculate the relative formula mass of the substance. (*1 mark*)

3 Calcium carbonate breaks down on heating to produce calcium oxide and carbon dioxide gas.

$$CaCO_3 \rightarrow CaO + CO_2$$

A student heats 15 g of calcium carbonate strongly in a crucible.

Relative atomic masses (A_r): Ca = 40, C = 12, O = 16.

a Calculate the number of moles in 15 g of calcium carbonate. (*2 marks*)

b Use your answer to part **a** to determine the number of moles of calcium oxide that will be produced. (*1 mark*)

c Calculate the mass of calcium oxide produced by this reaction. (*2 marks*)

4 A student is preparing a sample of sodium nitrate, $NaNO_3$.

She mixes 17 g of sodium oxide with an excess of nitric acid. The equation for the reaction is:

$$Na_2O + 2HNO_3 \rightarrow 2NaNO_3 + H_2O$$

Relative atomic masses (A_r): Na = 23, O = 16, N = 14, H = 1.

a Calculate the number of moles in 17 g of sodium oxide. (*2 marks*)

b Calculate the maximum mass of sodium nitrate that can be produced in this reaction. (*3 marks*)

5 Aluminium and iron oxide (Fe_2O_3) react together to produce aluminium oxide (Al_2O_3). The equation for the reaction is:

$$2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$$

Calculate the mass of iron that is produced by reacting 20 g of iron oxide with an excess of aluminium.

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Relative atomic masses (A_r): Al = 27, O = 16, Fe = 56. (*4 marks*)

Maths skills links

You may also need to convert the mass of a substance into an amount in moles (and vice versa) when using moles to balance equations.

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Name .... **Date** ....

Standard form and significant figures

Specification references

- *MS 0.1 Recognise and use expressions in standard and ordinary form*
- *MS 0.4 Use calculators to find and use power, exponential and logarithmic functions*
- *MS 1.1 Use an appropriate number of significant figures*

Learning objectives

After completing the worksheet you should be able to:

- convert between numbers in standard and ordinary form
- state numbers to a certain degree of accuracy.

Introduction

In the calculations you will be asked to perform as part of your AS studies you will need to be confident with both representing numbers in standard form and giving them to a certain number of significant figures.

When numbers are very large or very small they are written in **standard form**. In standard form a number is written in the format:

$$a \times 10^{n} \text{ where } 1 \leq a < 10 \text{ and } n \text{ is an integer.}$$

In an experiment, or from a calculation, you may only be able to give your answer with a certain amount of accuracy. This accuracy is shown by giving your answer to a certain number of **significant figures**.

Worked example: Standard form

Question

Express 0.00268 in standard form.

Answer

Step 1

Identify the value for '*a*.' In this case it will be 2.68.

Step 2

Work out how many places the decimal place must be moved to form this number.

0 . 0 0 2 6 8

The decimal place must move 3 places to the right to become 2.68.

This number of places is the value for the integer '*n*.' If the decimal point moves to the right '*n*' is negative. If the decimal place moves to the left '*n*' is positive.

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Step 3

Substitute your values into the general format, $a \times 10^n$

Therefore in standard form 0.00268 is 2.68×10^{-3}.

Worked example: Significant figures

Question

Express 0.56480900 to 3 significant figures.

Answer

Step 1

Identify the numbers which are significant using the rules below:

Rule 1 Any number that isn't 0 is significant.

Rule 2 Any 0 that is between two numbers that are not 0 is significant.

Rule 3 Any 0 that is before all the non-zero digits is not significant.

Rule 4 Any 0 that is after all of the non-zero digits is only significant if there is a decimal point.

In this case the significant numbers are 0.**564 809** 00.

Step 2

Identify the three most significant figures. These are the significant numbers which are furthest to the left (have the biggest values), i.e., 0.**564** 809 00.

Step 3

Look at the next number. If this number is 5 or above, then round up. If this number is 4 or less, do not round up.

In this case the next number is 8, so we round up to 0**.565**.

Questions

1 This question is about expressing numbers in standard form.

a Express the following numbers in standard form. (*4 marks*)

i 0.0023

ii 1032

iii 275 000 0

iv 0.000528

b Write out the following numbers in ordinary form. (*4 marks*)

i 2.01×10^3

ii 5.2×10^{-2}

iii 8.41×10^2

iv 1.00×10^{-4}

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c For each of the pairs of numbers below identify which is the bigger number. (*3 marks*)

v 1.43×10^{23} or 1.43×10^{24}

vi 5.16×10^{-3} or 5.16×10^{-4}

vii 12.4×10^{23} or 1.50×10^{24}

2 Express the following numbers to the number of significant figures indicated. (*6 marks*)

a 4.74861 to two significant figures

b 507980 to three significant figures

c 809972 to three significant figures

d 06.345 to three significant figures

e 7840 to three significant figures

f 0.007319 to three significant figures

3 Carry out the following calculations expressing the numbers in **standard form** to the degree of accuracy indicated: (*4 marks*)

a $(4.567 \times 10^{5}) \times (2.13 \times 10^{-3})$ to three significant figures

b $(1.567 \times 10^{3}) \div (2.245 \times 10^{-1})$ to four significant figures

c $(5.4 \times 10^{-1}) \div (2.7 \times 10^{-3})$ to one significant figure

d $(2.00 \times 10^{-2}) \times (2.00 \times 10^{-4})$ to three significant figures

Maths skills links to other areas

You will use these skills throughout the Amount of substance topics.